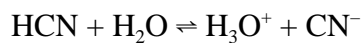
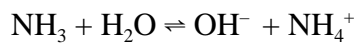


Weak Acids, Weak Bases and Buffers

- 1) Write the Brønsted-Lowry reaction for weak acid HCN reacting with H₂O.



- 2) Write the Brønsted-Lowry reaction for weak base NH₃ reacting with H₂O



- 3) Using the information from the Chemical Equilibria Exercises, write the equilibria expression for the acid HCN in 1) above using the substitution $K_a = K$

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{CN}^-]}{[\text{HCN}]}$$

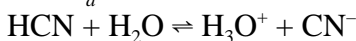
- 4) Write the equilibrium expression for the base NH₃ in 2) above using the substitution $K_b = K$.

$$K_b = \frac{[\text{OH}^-][\text{NH}_4^+]}{[\text{NH}_3]}$$

Weak Acids, Weak Bases and Buffers

5) Calculate the pH of a solution which is 0.010 M in HCN.

The K_a of HCN is 4.2×10^{-10} .



— x —>

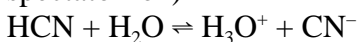
	[HCN]	[H ₃ O ⁺]	[CN ⁻]
initially	0.010	0	0
equilibrium	0.010 - x	x	x

substituting: $4.2 \times 10^{-10} = \frac{x^2}{0.010 - x}$ and assume $x \ll 0.010$

$x^2 = 4.2 \times 10^{-12}$ so: $x = 2.0 \times 10^{-6}$

Ans: 5.69

6) Calculate the pH of a solution which is 0.10 M in HCN and 0.10 M in NaCN. The K_a of HCN is 4.2×10^{-10} . (Notice that the important chemical species here is the CN⁻ ion. Na⁺ is a spectator ion)



— x —>

	[HCN]	[H ₃ O ⁺]	[CN ⁻]
initially	0.010	0	0
equilibrium	0.10 - x	x	0.10 + x

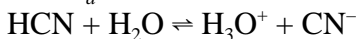
substituting: $4.2 \times 10^{-10} = \frac{(0.10 - x)x}{0.10 - x}$ and assume $x \ll 0.10$

$x = 4.2 \times 10^{-10}$

Ans: 9.38

7) Calculate the pH of a solution which is 0.10 M in HCN and 0.30 M in NaCN.

The K_a of HCN is 4.2×10^{-10} . (What is Na⁺? Ans: spectator ion)



— x —>

	[HCN]	[H ₃ O ⁺]	[CN ⁻]
initially	0.010	0	0
equilibrium	0.10 - x	x	0.30 + x

substituting: $4.2 \times 10^{-10} = \frac{(0.30 - x)x}{0.10 - x}$ and assume $x \ll 0.10$ (and $\ll 0.30$)

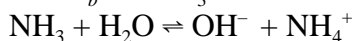
$x = 1.4 \times 10^{-10}$

Ans: 9.85

Weak Acids, Weak Bases and Buffers

8) Calculate the pH of a solution which is 0.010 M in NH_3 .

The K_b for NH_3 is 1.8×10^{-5} .



— x —>

	$[\text{NH}_3]$	$[\text{OH}^-]$	$[\text{NH}_4^+]$
initially	0.010	0	0
equilibrium	$0.010 - x$	x	x

substituting: $1.8 \times 10^{-5} = \frac{x^2}{0.010 - x}$ and assume $x \ll 0.010$

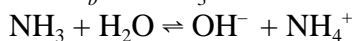
$x^2 = 1.8 \times 10^{-7}$ so: $x = 4.2 \times 10^{-4}$ ($=[\text{OH}^-]$)

$\text{pOH} = 3.37$

Ans: 10.63

9) Calculate the pH of a solution which is 0.010 M in NH_3 and 0.010 M in NH_4^+ .

The K_b for NH_3 is 1.8×10^{-5} .



— x —>

	$[\text{NH}_3]$	$[\text{OH}^-]$	$[\text{NH}_4^+]$
initially	0.010	0	0
equilibrium	$0.010 - x$	x	$0.010 + x$

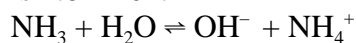
substituting: $1.8 \times 10^{-5} = \frac{(0.010 + x)x}{0.010 - x}$ and assume $x \ll 0.010$

$x = 1.8 \times 10^{-5}$ ($=[\text{OH}^-]$)

$\text{pOH} = 4.74$

Ans: 9.26

10) Calculate the *pH* of a solution which is 0.10 M in NH_3 and 0.010 M in NH_4^+ . The K_b for NH_3 is 1.8×10^{-5} .



— x —>

	$[\text{NH}_3]$	$[\text{OH}^-]$	$[\text{NH}_4^+]$
initially	0.10	0	0
equilibrium	$0.10 - x$	x	$0.010 + x$

substituting: $1.8 \times 10^{-5} = \frac{(0.010 + x)x}{0.10 - x}$ and assume $x \ll 0.010$ (and $\ll 0.10$)

$x = 1.8 \times 10^{-4}$ ($=[\text{OH}^-]$)

$\text{pOH} = 3.74$

Ans: 10.26

Problems 11, 12 and 13 are dilution problems. If you do not know how to do dilution problems, do some review first. These are in preparation for subsequent questions.

Weak Acids, Weak Bases and Buffers

- 11) If 100.0 mL of a solution is originally 0.10 M CH_3COOH and one adds 50.0 mL of water to this solution, what is the final concentration of the CH_3COOH ?

$$\text{For dilutions: } C_1V_1 = C_2V_2$$

$$C_1 = 0.10 \text{ M}$$

$$V_1 = 100.0 \text{ mL}$$

$$C_2 = ?$$

$$V_2 = 150 \text{ mL (volume of dilute solutions add with little error.)}$$

substitute:

$$(0.10 \text{ M})(100.0 \text{ mL}) = C_2(150 \text{ mL})$$

$$\text{Ans: } \underline{0.067 \text{ M}}$$

- 12) If 100.0 mL of a solution is originally 0.10 M CH_3COOH and one adds 50.0 mL of 0.10 M NaCl to this solution, what is the final concentration of the CH_3COOH ?

$$C_1V_1 = C_2V_2$$

For CH_3COOH :

(NaCl is irrelevant)

$$(0.10 \text{ M})(100.0 \text{ mL}) = C_2(150.0 \text{ mL})$$

$$\text{Ans: } \underline{0.067 \text{ M}}$$

- 13) If 100.0 mL of a solution is originally 0.10 M CH_3COOH and one adds 50.0 mL of 0.10 M NaCH_3COO to this solution, what is the final concentration of the CH_3COOH ? What is the final concentration of the NaCH_3COO ? (At this point, ignore the establishment of the Brønsted-Lowry equilibrium.)

$$C_1V_1 = C_2V_2$$

For CH_3COOH :

NaCH_3COO :

$$(0.10 \text{ M})(100.0 \text{ mL}) = C_2(150.0 \text{ mL}) \quad (0.10 \text{ M})(50.0 \text{ mL}) = C_2(150.0 \text{ mL})$$

$$\text{For M of } \text{CH}_3\text{COOH:} \quad \text{Ans: } \underline{0.067} \text{ M } \text{CH}_3\text{COOH}$$

$$\text{For M of } \text{NaCH}_3\text{COO} \quad \text{Ans: } \underline{0.033} \text{ M } \text{NaCH}_3\text{COO}$$

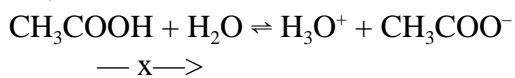
(What is Na^+ in NaCH_3COO ?) What is the important ion?

$$\text{Ans: the } \underline{\text{CH}_3\text{COO}^- \text{ (acetate)}} \text{ ion}$$

Weak Acids, Weak Bases and Buffers

- 14) What are the pHs of the solutions described in questions 11, 12 and 13? The K_a for $\text{CH}_3\text{COOH} = 1.8 \times 10^{-5}$.

For 11:



	$[\text{CH}_3\text{COOH}]$	$[\text{H}_3\text{O}^+]$	$[\text{CH}_3\text{COO}^-]$
initially	0.067	0	0
equilibrium	$0.067 - x$	x	x

substituting: $1.8 \times 10^{-5} = \frac{x^2}{0.067 - x}$ and assume $x \ll 0.10$ (and $\ll 0.30$)

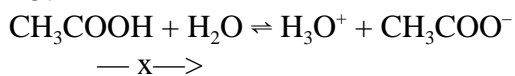
$x^2 = 1.2 \times 10^{-6}$ $x = 1.1 \times 10^{-3}$

Ans #11: pH = 2.96

For 12: Same as for 11 since Na^+ and Cl^- are a spectator ions.

Ans #12: pH = 2.96

For 13:



	$[\text{CH}_3\text{COOH}]$	$[\text{H}_3\text{O}^+]$	$[\text{CH}_3\text{COO}^-]$
initially	0.067	0	0.033
equilibrium	$0.067 - x$	x	$0.033 + x$

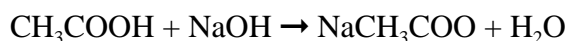
substituting: $1.8 \times 10^{-5} = \frac{(0.033 + x)x^2}{0.067 - x}$ and assume $x \ll 0.033$ (and $\ll 0.067$)

$x = 3.6 \times 10^{-5}$

Ans #13: pH = 4.44

- 15) A solution is created by mixing 100.0 mL of a solution which is 0.10 M CH_3COOH with 50.0 mL of a solution which is 0.10 M NaOH . (Confused? - Look up Arrhenius Acid/Base and neutralization.)

a) Write the Arrhenius acid-base reaction for this:



b) What is the concentration of the CH_3COOH if one does not take into account the Brønsted-Lowry reaction with water?

By definition $C = n/V$ where V is in liters.

and so $n = CV$

Weak Acids, Weak Bases and Buffers

moles of:	when:	How calculated:	ANS:
CH ₃ COOH	initially	(0.10 M)(0.1000 L)	0.010 mol
NaOH	initially	(0.10 M)(0.0500 L)	0.0050 mol
NaCH ₃ COOH	after Arrhenius reaction	$n_{\text{NaCH}_3\text{COO}}/1 = n_{\text{NaOH}}/1^*$	0.0050 mol
CH ₃ COOH	after Arrhenius reaction	$n_{\text{starting}} - n_{\text{used}}$	0.005 mol

*limiting reactant.

$$\text{thus: } C_{\text{CH}_3\text{COOH}} = \frac{0.0050 \text{ mol}}{0.1500 \text{ L}} \quad \text{Ans: } \underline{0.033} \text{ M CH}_3\text{COOH}$$

15 continued:

- c) What is the concentration of the NaCH₃COO if one does not take into account the Brønsted-Lowry reaction with water?

$$\text{thus: } C_{\text{NaCH}_3\text{COO}} = \frac{0.0050 \text{ mol}}{0.1500 \text{ L}} \quad \text{Ans: } \underline{0.033} \text{ M NaCH}_3\text{COO}$$

- d) What is the *pH* of the resultant solution described above?
The K_a for CH₃COOH = 1.8×10^{-5} .

$$\text{pH} = \text{p}K_a = -\log(1.8 \times 10^{-5}) = 4.74 \quad (\text{Why?})$$

$$\text{Ans: } \text{pH} = \underline{4.74}$$

- 16) What is the *pH* of a solution formed by mixing 200 mL of 0.20 M NH₃ with 50 mL of water? K_b for NH₃ = 1.8×10^{-5} .

$$C_1V_1 = C_2V_2$$

$$C_1 = 0.20 \text{ M}$$

$$V_1 = 200 \text{ mL}$$

$$C_2 = ?$$

$$V_2 = 250 \text{ mL}$$

$$(0.20 \text{ M})(200 \text{ mL}) = C_2(250 \text{ mL}) \quad C_2 = 0.16 \text{ M}$$

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

$$1.8 \times 10^{-5} = \frac{x^2}{0.16 - x}$$

$$x^2 = 2.9 \times 10^{-6} \quad x = 1.6 \times 10^{-3} (= [\text{OH}^-]) \quad \text{Ans: } \text{pH} = \underline{11.23}$$

Weak Acids, Weak Bases and Buffers

- 17) What is the pH of a solution formed by mixing 200 mL of 0.20 M NH_3 with 50 mL of 0.50 M NH_4Cl ? The K_b for $\text{NH}_3 = 1.8 \times 10^{-5}$.

$$C_1V_1 = C_2V_2 \text{ for both } [\text{NH}_3] \text{ and } \text{NH}_4^+$$

$[\text{NH}_3]$:

$$C_1 = 0.20 \text{ M}$$

$$V_1 = 200 \text{ mL}$$

$$C_2 = ?$$

$$V_2 = 250 \text{ mL}$$

$$C_2 = 0.16 \text{ M}$$

$[\text{NH}_4^+]$:

$$C_1 = 0.50 \text{ M}$$

$$V_1 = 50 \text{ mL}$$

$$C_2 = ?$$

$$V_2 = 250 \text{ mL}$$

$$C_2 = 0.10 \text{ M}$$

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

$$1.8 \times 10^{-5} = \frac{(0.10 + x)x}{0.16 - x}$$

$$x = 2.9 \times 10^{-5} (= [\text{OH}^-])$$

$$\text{pOH} = 4.54$$

$$\text{Ans: pH} = \underline{9.46}$$

- 18) What is the pH of a solution formed by mixing 200 mL of 0.20 M NH_3 with 50 mL of 0.50 M HCl ? (Note: you must first answer questions similar to question 17 above. What is the concentration of NH_3 and NH_4Cl ?) The K_b for $\text{NH}_3 = 1.8 \times 10^{-5}$.

Final volume = 250 mL

$$n_{\text{NH}_3} \text{ originally} = (0.20 \text{ M})(200 \text{ mL}) = 40 \text{ mmol}$$

$$n_{\text{HCl}} \text{ originally} = (0.50 \text{ M})(50 \text{ mL}) = 25 \text{ mmol (limiting reactant)}$$

$$n_{\text{NH}_4\text{Cl}} \text{ produced} = n_{\text{HCl}} \text{ originally} = 25 \text{ mmol}$$

$$n_{\text{NH}_3} \text{ left} = 40 \text{ mmol} - 25 \text{ mmol} = 15 \text{ mmol}$$

$$\text{resultant } [\text{NH}_3] = 15 \text{ mmol}/250 \text{ mL} = 0.060$$

$$\text{resultant } [\text{NH}_4^+] = 25 \text{ mmol}/250 \text{ mL} = 0.10$$

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

$$1.8 \times 10^{-5} = \frac{(0.10 + x)x}{0.060 - x}$$

$$x = 1.1 \times 10^{-5} (= [\text{OH}^-])$$

$$\text{pOH} = 4.97$$

$$\text{Ans: pH} = \underline{9.03}$$

- 19) What is the pH of a solution formed by mixing 200 mL of 0.20 M NH_3 with 50 mL of 0.50 M NH_4Cl and 50 mL of 0.50 M HCl ? (Hint: first calculate the dilutions for each of these species, then do the Arrhenius acid-base reaction.) The K_b for $\text{NH}_3 = 1.8 \times 10^{-5}$.

Final Volume = 300 mL

$$n_{\text{NH}_3} \text{ originally} = (0.20 \text{ M})(200 \text{ mL}) = 40 \text{ mmol}$$

$$n_{\text{HCl}} \text{ originally} = (0.50 \text{ M})(50 \text{ mL}) = 25 \text{ mmol (limiting reactant)}$$

$$n_{\text{NH}_4\text{Cl}} \text{ originally} = (0.50 \text{ M})(50 \text{ mL}) = 25 \text{ mmol}$$

$$n_{\text{NH}_4\text{Cl}} \text{ produced} = n_{\text{HCl}} \text{ originally} = 25 \text{ mmol}$$

$$n_{\text{NH}_4\text{Cl}} \text{ total} = 25 \text{ mmol} + 25 \text{ mmol} = 50 \text{ mmol}$$

$$n_{\text{NH}_3} \text{ left} = 40 \text{ mmol} - 25 \text{ mmol} = 15 \text{ mmol}$$

$$\text{resultant } [\text{NH}_3] = 15 \text{ mmol}/300 \text{ mL} = 0.050$$

$$\text{resultant } [\text{NH}_4^+] = 50 \text{ mmol}/300 \text{ mL} = 0.167$$

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]}$$

$$1.8 \times 10^{-5} = \frac{(0.167 + x)x}{0.050 - x}$$

$$x = 5.4 \times 10^{-6} (= [\text{OH}^-])$$

$$\text{pOH} = 5.27$$

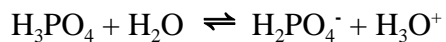
$$\text{Ans: pH} = \underline{8.73}$$

Weak Acids, Weak Bases and Buffers

Use the following K_a s to answer questions 20 through 23:

For the reaction:

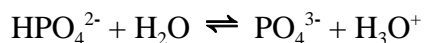
The equilibrium constant is:



$$K_{a1}(\text{H}_3\text{PO}_4) = 7.5 \times 10^{-3}$$



$$K_{a2}(\text{H}_2\text{PO}_4^-) = 6.2 \times 10^{-8}$$



$$K_{a3}(\text{HPO}_4^{2-}) = 3.6 \times 10^{-13}$$

20) What is the pH of a solution which is 0.10 M in H_3PO_4 ?

$$K_{a1} = \frac{[\text{H}_3\text{PO}_4][\text{H}_3\text{O}^+]}{[\text{H}_2\text{PO}_4^-]}$$

$$7.5 \times 10^{-3} = \frac{x^2}{0.10 - x}$$

etc.

Ans: $pH = \underline{1.56}$ (1.62 if $x \ll 0.10$ not assumed)

21) What is the pH of a solution which is 0.10 M in H_3PO_4 and 0.10 M in NaH_2PO_4 ?

(What is Na^+ ? It is spectator ion)

$$pH = pK_{a1}$$

Ans: $pH = \underline{2.12}$

22) What is the pH of a solution which is 0.10 M in NaH_2PO_4 and 0.10 M in Na_2HPO_4 ?

$$pH = pK_{a2}$$

Ans: $pH = \underline{7.21}$

23) What is the pH of a solution which is 0.10 M in Na_2HPO_4 and 0.10 M in Na_3PO_4 ?

$$pH = pK_{a3}$$

Ans: $pH = \underline{12.44}$